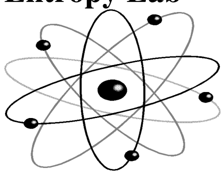

Entropy Lab! (A very disorderly lab)

Introduction:

Reactions in nature are driven by two forces, which in combination determine whether or not the reaction will be spontaneous. Firstly, reactions that are exothermic (give off heat) are generally favored by nature. However, some endothermic reactions, such as the melting of ice, are spontaneous and thus favored by nature. The second driving force that also determines whether or not a reaction will occur is entropy. Entropy can be defined as a measure of the degree of randomness of the particles, such as molecules, in a system. As might be expected from the chaotic world in which we live, nature favors an increase in entropy. In other words, reactions that increase the disorder of the system tend to be spontaneous. The amount of entropy of a system is best understood by considering the three principle states of matter. In a solid, the particles vibrate in place and are not free to switch places with each other. As such, solids are considered to have very low entropy because very little randomness exists in them.[1] Liquids, in general, are more disorderly than solids, and thus have higher entropy. Gases, the most disorderly of the three states possesses the highest entropy. These are general guideline as some liquids (mercury, for instance) have lower entropy that certain solids. In general, the dissolution process increases the entropy of a system. Entropy of substances can be determined quantitatively in the lab and are measured in molar values with units kJ/(mol·K). In this class, we will be most concerned with the change in entropy of a system, denoted ΔS.

Gibbs Free Energy

Reactions in nature tend toward decreasing enthalpy and increasing entropy. As such, a reaction that is exothermic and increases entropy will always be spontaneous. Conversely, an endothermic reaction that decreases the randomness of the system will never be spontaneous. However, what about reactions that decrease enthalpy but decrease entropy, or increase entropy but increase enthalpy? Will these reaction be spontaneous? The answer is that it depends on the temperature. In order to predict this, we must look at a factor called Gibbs Free Energy. Mathematically, Gibbs free energy relates the enthalpy and entropy changes of a reaction:

$$\Delta G_0 = \Delta H_0{}^{-}\ T\Delta S_0$$

If the Gibbs Free Energy of a reaction is negative, then the reaction will be spontaneous, if positive, then the reaction will not be spontaneous.

ΔH	**ΔS**	**ΔG**
- value (exothermic)	+ value (more random)	Always negative
- value (exothermic)	- value (less random)	Negative at lower temperature
+ value (endothermic)	+ value (more random)	Negative at higher temperature
+ value (endothermic)	- value (less random)	Never negative

[1] The entropy of a perfect crystalline solid at absolute zero (0 K; -273.15 ^{0}C) is defined to be zero. As energy is added molecular motion and randomness increase.

Purpose: To study reactions and determine the driving forces behind the reactions.

Materials:

1. Sodium Bicarbonate
2. Acetic Acid
3. Glass Test Tube
4. Wood Splint
5. Matches
6. Ethanol
7. Ammonium Chloride
8. Thermometer
9. Potassium permanganate
10. Concentrated Sulfuric Acid

Procedure:

<u>Test #1</u>

1. Measure out 10 ml of 1.0 M sodium bicarbonate into a test tube. Carefully add 10 ml of 3 M acetic acid to the test tube.
2. Record your observations of the reaction.
3. Light a wood splint using matches. Blow out the flame and place the glowing splint in the mouth of the test tube. (Be careful not to drop the splint into the test tube).
4. Record your observations
5. Dispose of your solution as directed by your teacher and clean your glassware before proceeding.

<u>Test #2</u>

1. Add 10 ml of ethanol to a clean small plastic bottle.
2. Swirl the contents of the bottle for approximately 2 minutes, and drain the ethanol down the drain.
3. Let the bottle stand for 30 seconds. Light a wood splint with a match and carefully pass it over the mouth of the bottle.
4. Record your observations.

<u>Test #3</u>

1. Measure out 100 ml of distilled water in a 250 ml beaker and record its temperature.
2. Measure out 2.50 grams of ammonium chloride. Add the ammonium chloride to the distilled water.
3. Record the temperature of the solution after the dissolution process is complete.

<u>Test #4</u>

1. Measure out a 3 cm length of copper wire. Place the copper wire in a 250 ml beaker.
2. In a fume hood measure out 20 ml of 6.0 M nitric acid.
3. Carefully add nitric acid to beaker in fume hood.
4. Record your observations.
5. Place watch glass over beaker, and allow reaction to proceed overnight.

Questions:

Test #1

1. What gas was produced in this test? How do you know?
2. Write out two complete balanced equations for the reactions. (The first reaction should be a double displacement, and the second reaction should be a decomposition of carbonic acid).
3. Were the reactions exothermic or endothermic?
   a. How might you change the experiment to determine this?
4. Based on your chemical equation, did the system's entropy increase or decrease in each reaction. Explain your reasoning?
5. Was the ΔG for the reaction positive or negative? How do you know?

Test #2

6. Write a complete balanced equation for the combustion of ethanol. (The products of this combustion are CO_2 (g) and H_2O (l)).
7. Is this product exothermic or endothermic? How can you tell?
8. Based on the number of gas molecules on each side of the equation, did the entropy of the system increase or decrease?
9. Was the ΔG for the reaction positive or negative? How do you know?
10. How might changing the temperature of the system affect the ΔG?
11. Consider the reaction 3 C (graphite) + 4 H_2 (g) $\rightarrow$ C_3H_8 (g) at 25.00 ºC
    a. [ΔH_{rxtn} = -103.8 kJ, ΔS_{rxtn} = -269.1 J/K]. Calculate the change in free energy for this reaction.

Test #3

12. Write out a complete balanced equation for this dissociation.
13. Was this reaction exothermic or endothermic? Explain your reasoning.
14. Did entropy of the system increase or decrease in this reaction? Explain your reasoning.
15. Was the ΔG for the reaction positive or negative? How do you know?
    a. What was the driving force of this reaction?
16. How might changing the temperature of the system affect the ΔG?

Test #4

17. Write out a complete balanced equation for the reaction between the nitric acid and the copper. (The products of the reaction are $(CuNO_3)_2$ (aq), H_2O, and NO_2 (g)).
18. Has the entropy of the system increased or decreased? Explain your reasoning.